ANSWERS: Ionisation Energy

1) First ionisation energy is the minimum energy required to remove one mole of electrons from one mole of gaseous atoms.
First ionisation energy increases from 502 in Na to1527 in Ar. There is an increase in the number of protons and thus the nuclear charge / attractive force of the nucleus. As the electrons are added to the same energy level, there is no increase in repulsion between energy levels. The nuclei with a greater number of protons have a stronger electrostatic attraction for the valence electrons in the third shell, thus the first ionisation energy increases across a period.
Both periodic trends are influenced by nuclear charge and the number of shells / distance, the ionisation energy increases while the atomic radii decrease.
The larger the ionisation energy the more strongly the valence electrons are held. Thus atomic radii across Period 3 decrease.

2) *lowest* B N Ne He *highest*

3) Cl has more protons than Li. Therefore there is a greater attraction between the nucleus and outer electrons/electrons held more tightly so it is harder to remove an electron from Cl than Li.
Even though the valence electrons of Cl are in the 3rd energy level/has an extra energy level the extra shielding is not as significant as the effect of the increased nuclear charge, so Cl has a higher first ionisation energy than Li.

4) Cl and Na have the same number of energy levels. However, Cl has a greater number of protons causing a stronger attraction to electrons than Na. Hence Cl has a greater first ionisation energy.

5) i) $Li(g) \rightarrow Li^{+}(g) + e^{-}$
ii) As you move across a period from Li to Ne, the ionisation energies increase. Electrons are added to the same valence shell / the same distance from the nucleus. Extra protons in the nucleus increase the nuclear charge, so the electrons in the valence shell are held more tightly and ionisation energy is greater.
As you go down a group, ionisation energy decreases. This is due to a new energy level being added, which is further from the nucleus. Electrons can be removed more easily and the ionisation energy is less
The drop Be and B is due to B having 1 electron in the p subshell ($2p^1$) and Be being $2s^2$. Although B has a greater nuclear charge, the electron in the p-subshell is further from the nucleus/has less stability. Thus the p-electron in B's valence shell is not held so tightly/is more easily removed.
Drop N – O N has 1/2 full subshell ($2p^3$) and O 1 more electron giving it a partly full subshell ($2p^4$). added electron going into suborbital already occupied by an electron – increased electron electron repulsion so makes electron more easily removed / partly full subshell less stable so electron more easily removed

6) Valence electrons are added to same shell / distance from nucleus similar.
In Br, there is a greater number of protons / nuclear attraction greater, so <u>valence electron more strongly held</u> (implying IE).

7) Br electrons closer to the nucleus / smaller radius / bromine is smaller, Br has greater nuclear charge / number of protons, but same number of shells / energy levels
Br has greater nuclear charge than Sc, causing stronger attraction to the electrons

8) i) Energy required to remove the outermost electron from one mole of gaseous atoms $Cl(g) \rightarrow Cl^{+}(g) + e^{-}$
ii) Ca < Mg < Cl
Ionisation energies increase across the table or decrease down the group.
Ca has the greatest number of electron shells, so its electrons are further from the nucleus, so less energy is required to remove a valence electron. Ca also has greater shielding effect of an additional shell between the valence shell and the nucleus, so less energy is required to remove the outermost electron.
Mg and Cl have same number of shells, but Cl has greater nuclear charge or number of protons in the nucleus, so greater electrostatic attraction between nucleus and valence electrons.